

3A:  $\text{HCl}, \text{HI}, \text{HBr}$   
 $\text{HNO}_3, \text{H}_2\text{SO}_4, \text{HClO}_4$

Quiz 4, Ch 15  
 Due Tues 7/19

Name: KEY

3 points each:

1. Which of the following neutralization reactions creates a basic solution? (More than one MAY be correct: circle all that apply)  $\rightarrow$  SB, WA or WB/WA(?)

- a.  $\text{KOH}$  and  $\text{HNO}_3$   
 b.  $\text{HClO}_4$  and  $\text{N}_2\text{H}_4$   
 c.  $\text{HF}$  and  $\text{Ba}(\text{OH})_2$   
 d.  $\text{HCN}$  and  $\text{C}_2\text{H}_5\text{N}$

2. Which of the following gives a buffer solution when 50 mL each of the two solutions are mixed?

- a. 0.10 M  $\text{NH}_3$  and 0.10 M  $\text{HCl}$   
 b. 0.20 M  $\text{NH}_3$  and 0.10 M  $\text{HCl}$   
 c. 0.10 M  $\text{NH}_4\text{Cl}$  and 0.10 M  $\text{NH}_3$   
 d. 0.20 M  $\text{NH}_4\text{Cl}$  and 0.10 M  $\text{NaOH}$

buffer = weak acid & a  
 salt of it's  
 conjugate base  
 (or WB & it's conja)

3. What would the pH of the buffer solution in #2 be?

$$\text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{pK}_a = -\log \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 9.26$$

$$\text{pH} = 9.26 + \log \frac{(0.10)}{(0.10)} = 9.26$$

( $K_b = 1.8 \times 10^{-5}$   $K_w = 1 \times 10^{-14}$   $K_a = K_w / K_b$ )

4. What indicator (table 14.4) would be appropriate for the titration of 50 mL of 0.10 M  $\text{H}_2\text{CO}_3$  with 0.20 M  $\text{NaOH}$ ?

need to find pH at eq pt: need to find eq pt

$$0.10 \text{ mol/L} \times 0.050 \text{ L} \text{ H}_2\text{CO}_3 = 0.005 \text{ mol H}_2\text{CO}_3 \text{ BUT } 2 \text{ mol H}_3\text{O}^+ \text{ per H}_2\text{CO}_3$$

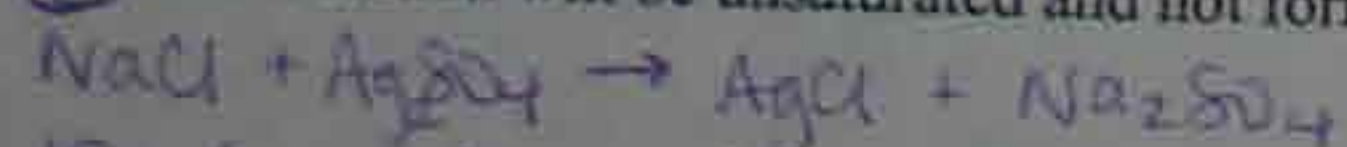
so 0.010 mol  $[\text{H}_3\text{O}^+]$  need 0.010 mol base

$$0.010 \text{ mol} \times \frac{\text{L}}{0.20 \text{ mol}} \text{ NaOH} = 0.05 \text{ L} = 50 \text{ mL base}$$

salt driven pH:  $\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{HCO}_3^- + \text{H}_3\text{O}^+$  see next pg.

5. 10 mL of a 0.01 M solution of sodium chloride is added to 40 mL of a 0.02 M solution of silver sulfate. What will happen?

- a. The solution will be supersaturated and form a precipitate  
 b. The solution will be saturated and therefore at equilibrium  
 c. The solution will be unsaturated and not form a precipitate



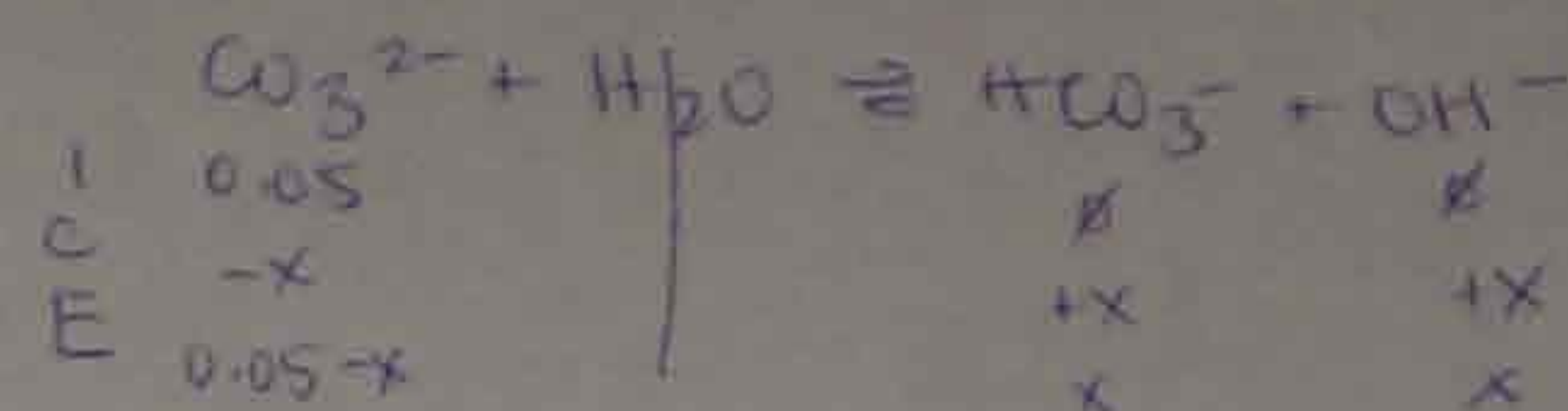
$$\text{IP} = (0.002)(0.032) = 6.4 \times 10^{-5} < K_{sp} (1 \times 10^{-4})$$

new total volume =  
 50 mL acid + 50 mL base =  
 100 mL

$$M_{\text{Cl}} = \frac{0.01(10 \text{ mL})}{50 \text{ mL}} = 0.002$$

$$M_{\text{Ag}} = \frac{2(0.02 \text{ M})(40 \text{ mL})}{50 \text{ mL}} = 0.032$$

2 Ag per molecule  
 $\therefore$  no ppt



$$\frac{0.10 \text{ M (50 mL)}}{100 \text{ mL}} = 0.05 \text{ M}$$

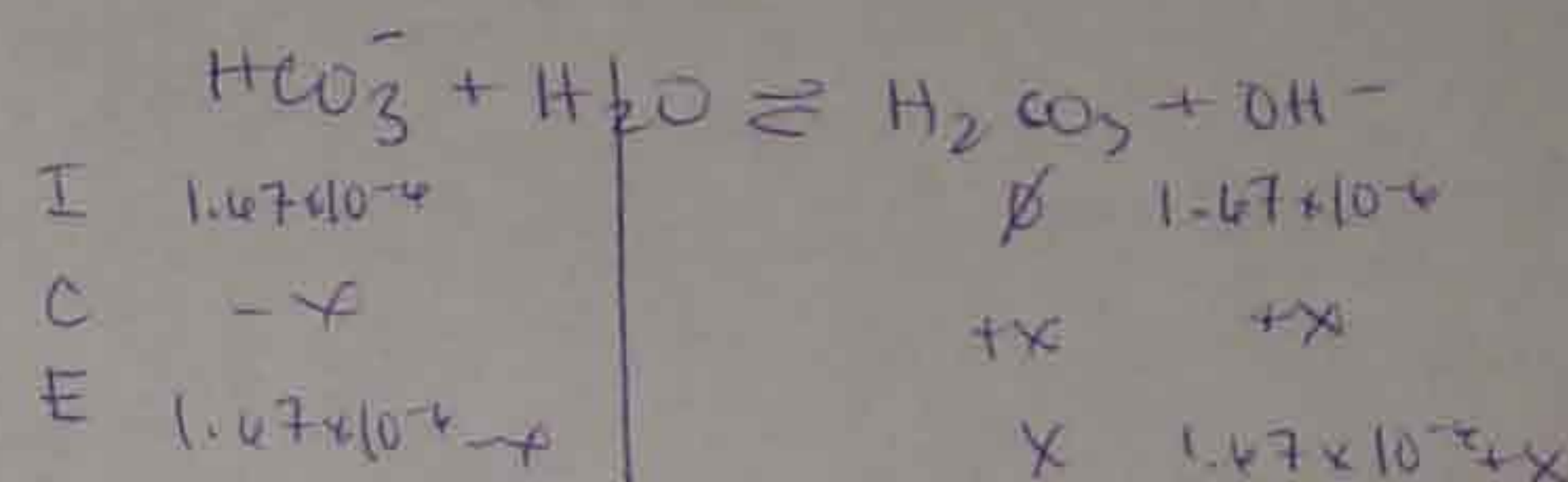
$$5.6 \times 10^{-11} = \frac{x^2}{0.05-x}$$

$\uparrow$  negligible

$$x^2 = (0.05)(5.6 \times 10^{-11})$$

$$x = \sqrt{\quad}$$

$$x = 1.67 \times 10^{-6}$$



$$4.3 \times 10^{-7} = \frac{x(1.67 \times 10^{-4} + x)}{(1.67 \times 10^{-4} - x)}$$

$$4.3 \times 10^{-7} (1.67 \times 10^{-4} - x) = 1.67 \times 10^{-4} x + x^2$$

$$7.18 \times 10^{-13} - 4.3 \times 10^{-7} x = 1.67 \times 10^{-4} x + x^2$$

$$0 = x^2 + 2.1 \times 10^{-4} x - 7.18 \times 10^{-13}$$

$$x = \frac{-2.1 \times 10^{-4} \pm \sqrt{(2.1 \times 10^{-4})^2 - 4(1)(-7.18 \times 10^{-13})}}{2}$$

$$x = \frac{-2.1 \times 10^{-4} \pm 2.7 \times 10^{-4}}{2}$$

$$x = \frac{0.6 \times 10^{-4}}{2} = 3 \times 10^{-7}$$

$$[\text{OH}^-] = 1.67 \times 10^{-4} + 3 \times 10^{-7} = 1.97 \times 10^{-4}$$

$$\text{pOH} = -\log(1.97 \times 10^{-4}) = 3.7$$

$$\text{pH} = 14 - \text{pOH} = 10.3$$